Stoichiometry
Volume-Volume Examples
(volume of gas, not solution)

Probs #1 - 10

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Brief Introduction: The key point to look for are the conditions of temperature and pressure. If they remain constant, you may treat the volumes in the same manner you treat moles. This is because, under conditions of constant T and P, the volumes are directly proportional to the moles. This is discussed in several of the problem solutions below.

While the great majority of volume-based stoichiometry problems are phrased in terms of constant T and P, they do not have to be. You can see this in problems 1b and 9, just below. If you have changing conditions of T and P, you, in general, will do this:

(1) Convert volme to moles using PV = nRT and the initial set of T and P
(2) Use a ratio and proportion to determine moles of other substance involved in problem
(3) Use PV = nRT with new T and P as well as moles of substance from step 2. You will calculate a new volume.

Be prepared! Your teacher could teach the situation where T and P do not change, then test on the situation where T and P do change.

Example #1: The equation for the combustion of methane is:

$$CH_4 + 2O_2 \text{ ---> } 2H_2O + CO_2$$

(a) If 50.0 L of methane at STP are burned, what volume of carbon dioxide will be produced at STP?
(b) If 50.0 L of methane at RTP are burned, what volume of gaseous water at STP is produced?

Solution to (a):

Because everything occurs at STP, the volumes are directly proportional to the moles of reactant used and product produced. Why? Consider PV = nRT. Rewrite it as:

n / V = P / RT

Everything on the right side is constant, so the n:V ratio must also be constant. That means that volume is directly proportional to the number of moles of gas.

Since there is a 1:1 molar ratio between CH_4 and CO_2, the answer is 50.0 L

Solution to (b):

1) Use PV = nRT to determine moles of methane:

(1.00 atm) (50.0 L) = (n) (0.08206 L atm mol^{-1} K^{-1}) (298 K)

n = 2.044665 mol

Note: in this problem, I am taking room temperature to be 25.0 °C since RTP means room temperature and pressure.

2) Use methane:water molar ratio:

1 is to 2 as 2.044665 mol is to x

x = 4.08933 mol of water produced

3) Use PV = nRT to calculate volume of water vapor at STP:

(1.00 atm) (V) = (4.08933 mol) (0.08206 L atm mol^{-1} K^{-1}) (273 K)

V = 91.6 L (to three sig figs)

Example #2: Given the following equation:

$C(s) + 2H_2(g) \rightarrow CH_4(g)$

How many liters of hydrogen are needed to produce 20.0 L of methane?

Solution:

1) P and T are not part of this problem:

This is because there is no mention whatsoever of T or P in the problem. We must assume them to be constant. If we assume they did change, but we don't know the values, the problem cannot be solved.

2) State Avogadro's Hypothesis (using moles rather than molecules):

Equal volumes of gas at equal temperature and pressure contain equal number of moles

3) State the hydrogen-methane molar ratio:

2:1

4) Use a ratio and proportion:

2 moles hydrogen is to 1 mole methane as x liters of hydrogen is to 20.0 L of methane

x = 40.0 L of hydrogen

Example #3: 2.35 L of oxygen gas reacts with 3.72 L of hydrogen gas, forming water. How many liters of the excess reactant will remain? If 2.50 L of water were actually produced, what would be the percent yield?

Solution:

1) The balanced chemical equation is:

$2H_2 + O_2 \rightarrow 2H_2O$

2) Determine the limiting reagent:

oxygen: 2.35 / 1 = 2.35
hydrogen: 3.72 / 2 = 1.86

Hydrogen is the limiting reagent.

Please note that no mention of temperature or pressue is made in the problem. This means that everything takes place at an unchanged temperature and pressure. Consequently, those two values remain constant and drop out of consideration. We do not need moles because, in a situation of constant temperature and pressure, the volumes are directly proportional to the number of moles.

Besides which, we cannot even calculate moles since we do not know the temperature or the pressure.

3) Use H_2:O_2 molar ratio:

2 is to 1 as 3.72 is to x

1.86 L of oxygen are used.

2.35 − 1.86 = 0.49 L of unreacted oxygen remain

4) Percent yield:

3.72 L of water is produced.

2.50 / 3.72 = 67.2%

Example #4: How much air is needed (in m^3, at 25.0 °C, 1.00 atm) to completely burn 10.0 moles of propane (C_3H_8). Assume that the air is composed of 21.0% O_2.

Solution:

1) The combustion of propane:

$C_3H_8 + 5O_2 \text{ ---> } 3CO_2 + 4H_2O$

2) Determine moles of pure oxygen needed to burn 10.0 mol of propane:

1 is to 5 as 10.0 is to x

x = 50.0 mol of oxygen required

3) Use PV = nRT to convert mol of oxygen to liters of oxygen at the conditions stated in the problem:

(1.00 atm) (V) = (50.0 mol) (0.08206 L atm mol^{-1} K^{-1}) (298 K)

V = 1222.694 L

4) Convert to volume of air required:

1222.694 L / 0.21 = 5822.35 L

5) Convert to cubic meters:

5822.35 L = 5822.35 dm^3

$1\ m^3 = 1000\ dm^3$

$(5822.35\ dm^3)\ (1\ m^3 / 1000\ dm^3) = 5.82\ m^3$ (to three sig figs)

Example #5: Propane (C_3H_8) burns in oxygen to produce carbon dioxide and water vapor. What volume of carbon dioxide is produced when 2.80 liters of oxygen are consumed.

Solution:

Since no T or P is given, we assume the reaction happens at constant temp and press.

That means that the coefficients of the balanced equation represent the volumetric ratio that the substances react in.

5 volumes of O_2 will produce 3 volumes of CO_2

From that, we find:

5/3 equals 2.8/x

x = 1.68 L

If we assumed the T and P changed in the reaction, we would not be able to solve the problem. However, do not make this assumption. Your teacher would be displeased.

Example #6: Methane burns according to the following equation:

$CH_4 + 2O_2 \text{ ---> } CO_2 + 2H_2O$

Calculate the volume of air that is required to burn 10.0 L of methane when both are at the same temperature and pressure. Assume that air is 20.0 percent oxygen by volume.

Solution:

The coefficients give the molar ratio that methane and oxygen react in when at the same T and P.

1 is to 2 as 10.0 is to x

x = 20.0 L of pure oxygen required.

20.0 L / 0.20 = 100.0 L of air is required.

Example #7: Nitrogen monoxide reacts with oxygen according to the equation below:

$2NO(g) + O_2(g) \text{ ---> } NO_2(g)$

How many liters of NO (reacting with excess oxygen) are required to produce 3.0 liters of NO_2?

Solution:

1) State pertinent ratio of volumes:

2:1

2) Write a ratio and proportion:

2 is to 1 as x is to 3.0 L

x = 6.0 L of NO required

By now, you should know what no mention of temperature or pressure means for solving the problem. (Hint: it means they remain constant throughout the entire problem, so they drop out of consideration.)

Example #8: If stoichiometric quantities of X and Y are placed in a sealed flexible container with an initial volume of 30.0 L at STP, what volume of Z will be produced? (X, Y, and Z, are all present in the gaseous state)

2X + Y ---> Z

Solution:

Seeing as they are all gases, a mole ratio is equal to the volume ratio:

2X + Y = 30.0 L

X : Y = 2 : 1 = 20L : 10L

n(Y) = n(Z)

therefore, v(Y) = v(Z) = 10.0 L of Z produced.

Example #9: 200.0 liters of H_2 reacting at 25.0 °C and 751.0 torr will require how many liters of O_2 at STP?

Solution:

1) The equation is:

$2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$

2) Use PV = nRT to calculate moles of hydrogen:

(751.0 torr / 760.0 torr/atm) (200.0 L = (n) (0.08206 L atm mol^{-1} K^{-1}) (298 K)

n = 8.0818 mol

3) Use H_2 : O_2 molar ratio:

2 is to 1 as 8.0818 ml is to x

x = 4.0409 mol of O_2 required

4) Use PV = nRT to calculate volume of O_2 at SPT that is required:

(1.00 atm) (V) = (4.0409 mol) (0.08206 L atm mol^{-1} K^{-1}) (273 K)

V = 90.52 L

Example #10: A mixture is prepared using 12.0 L of NH_3 and 12.0 L of Cl_2, both measured at the same conditions. These substances react according to the following equation:

$2\ NH_3(g) + 3\ Cl_2(g) \text{---> } N_2(g) + 6\ HCl(g)$

When the reaction is completed, what is the volume of each gas (NH_3, Cl_2, N_2 and HCl, respectively)? Assume the final volumes are measured under identical conditions.

Solution:

1) We need a limiting reagent:

ammonia: 12 / 2 = 6
chlorine: 12 / 3 = 4

We have a winner! Chlorine.

Remember, I can treat the volumes in the same way I would moles. This is because, at constant T and P, the volumes are directly proportional to the number of moles.

2) Final volume of the ammonia:

2 is to 3 as x is to 12

x = 8.0 L of ammonia used

volume remaining is 12.0 L − 8.0 L = 4.0 L

3) Final volume of chlorine:

Zero!

4) Final volume of nitrogen:

3 is to 1 as 12 is to x

x = 4.0 L

5) Final volume of HCl:

3 is to 6 as 12 is to x

x = 24.0 L

Bonus Example: How many liters of O_2 are produced (at STP) when sufficient KO_2 reacts with 75.0 L of carbon dioxide at −25.0 °C and 215 kPa?

$4KO_2(s) + 2CO_2(g) \text{---> } 2K_2CO_3(s) + 3O_2(g)$

Solution #1:

1) Determine how many moles of CO_2 are present:

$PV = nRT$

$(215\ kPa)\ (75.0\ L) = (n)\ (8.31447\ L\ kPa / mol\ K)\ (248\ K)$

$n = 7.820121\ mol$

2) Determine moles of O_2 produced:

$$\frac{2}{3} = \frac{7.820121\ mol}{x}$$

$x = 11.73018\ mol$

3) Determine volume at STP:

$PV = nRT$

$(101.325\ kPa)\ (V) = (11.73018\ mol)\ (8.31447\ L\ kPa / mol\ K)\ (273\ K)$

$V = 263\ L$ (to three sig figs)

Solution #2:

1) Determine volume of CO_2 at STP (using the combined gas law):

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

$$\frac{(215\ kPa)\ (75.0\ L)}{248\ K} = \frac{(101.325\ kPa)\ (V_2)}{273\ K}$$

$V_2 = 175.1838\ L$

2) Determine volume of O_2 produced at STP:

$$\frac{2}{3} = \frac{175.1838\ L}{x}$$

$x = 263\ L$

Solution #3:

1) Convert liters of CO_2 to liters of O_2 produced ar the same temp. and press. as the CO_2:

$$\frac{2}{3} = \frac{75.0\ L}{x}$$

x = 112.5 L

2) Convert to STP (using the combined gas law):

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

$$\frac{(215\ \text{kPa})\ (112.5\ \text{L})}{248\ \text{K}} = \frac{(101.325\ \text{kPa})\ (V_2)}{273\ \text{K}}$$

$V_2 = 263$ L

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